Periodicity Name: _____

AP Chemistry Lecture Outline

valence orbitals: outer-shell orbitals

-- elements in the same group have the same valence-shell electron configuration

-- since valence e^- are involved in bonding, elements within a group have many of the same properties

Development of the Periodic Table

-- few elements appear in elemental form in nature (Au, Ag, Hg, a few others)

-- most are in combined forms with other elements

-- In 19th century, advances in chemistry allowed more elements to be identified.

1869: Independently, Dmitri Mendeleev (Russia) and Lothar Meyer (Germany) published classification schemes based on similarities in element properties.

** Mendeleev used his scheme to predict the existence of undiscovered elements, and so is given credit for inventing the first periodic table.

** He organized the elements by increasing atomic mass.

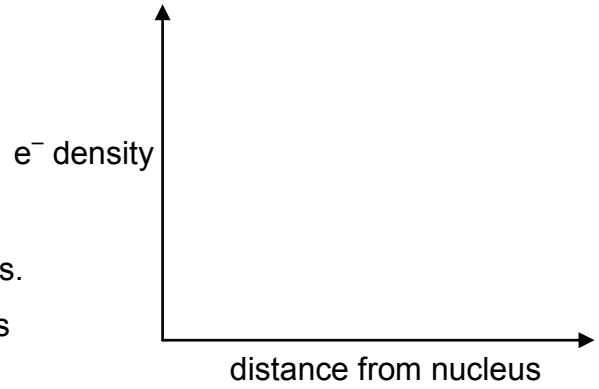
-- 1913: Henry Moseley bombarded atoms with high-energy electrons and measured the frequency of the X rays given off. X ray frequency generally increased as atomic mass increased, but not quite. Moseley rearranged the elements by atomic number.

Electron Shells

Even before Bohr, the American Gilbert Lewis had suggested that e^- are arranged in shells.

-- Experiments show that e^- density is a maximum at certain distances from nucleus.

-- no clearly defined boundaries between shells



Approximate bonding atomic radii for the elements have been tabulated.
The distance between bonded nuclei can be approximated by adding radii from both atoms.

e.g., Bonding atomic radii are as follows:

C = 0.77 A, Cl = 0.99 A

So approximate distance between bonded C and Cl nuclei =

Atomic Radius

As we go down a group, atomic radius…

--

As we go from left to right across the Table, atomic radius…

-- effective nuclear charge increases, but principal quantum number is constant

EX. Arrange the following atoms in order of increasing atomic radius: Sr, Ba, Cs

Ionization Energy: the minimum energy needed to remove an e^- from an atom or ion

$M(g) + 1^{st}$ I.E. →

$M^{1+}(g) + 2^{nd}$ I.E. →

$M^{2+}(g) + 3^{rd}$ I.E. →

Successive ionization energies are larger than previous ones.

-- (+) attractive force remains the same, but there is less electron-electron repulsion

The ionization energy increases sharply when we try to remove an inner-shell electron.

e.g.,

As we go down a group, first ionization energy…

--

Generally, as we go from left to right, first ionization energy…

Exceptions: e.g., B < Be Be: $1s^2\ 2s^2$

B: $1s^2\ 2s^2\ 2p^1$

e.g., O < N N: $1s^2\ 2s^2\ 2p^3$

O: $1s^2\ 2s^2\ 2p^4$

Electron affinity: the energy change that occurs when an e^- is added to a gaseous atom

For most atoms, adding an e^- causes energy to be…

Exceptions: noble gases: the added e^- must go into a new, higher energy level

group 2 metals: the added e^- must go into a higher-energy p orbital

group 15 elements: the added e^- is the first one to double-up a p orbital

The halogens have the most (–) electron affinities, meaning that they become very stable when they accept electrons.

Electron affinities don't vary much going down a group.

Regions of the Table

metals: left side of Table; form cations

properties:

-- Because of their low ionization energies, they are often oxidized in reactions.

-- Metallic character of the elements increases as we go down-and-to-the-left.

nonmetals: right side of Table; form anions

properties:

-- memorize the H O Br F I N Cl twins

metalloids (semimetals): "stair" between metals and nonmetals

properties:

Si and Ge →

Reactivity Trends

metal oxide + water ⟶ metal hydroxide

$Na_2O(s) + H_2O(l) \longrightarrow$

$CaO(s) + H_2O(l) \longrightarrow$

metal oxide + acid ⟶ salt + water

$MgO(s) + 2\ HCl(aq) \longrightarrow$

$NiO(s) + H_2SO_4(aq) \longrightarrow$

metal + nonmetal ⟶ salt

$2\ Al(s) + 3\ Br_2(l) \longrightarrow$

$2\ Fe(s) + O_2(g) \longrightarrow$

nonmetal oxide + water ⟶ acid

$CO_2(g) + H_2O(l) \longrightarrow$

$P_4O_{10}(s) + 6\ H_2O(l) \longrightarrow$

nonmetal oxide + base ⟶ salt + water

$CO_2(g) + 2\ NaOH(aq) \longrightarrow$

$SO_3(g) + 2\ KOH(aq) \longrightarrow$

Group Trends

Alkali Metals -- the most reactive metals (one e^- to lose)

-- obtained by electrolysis of a molten salt

e.g., chloride ion is oxidized and sodium ion is reduced

-- react with hydrogen to form metal hydrides:

-- react with water to form metal hydroxides:

-- react w/O_2: Li yields Li_2O, others yield peroxides (M_2O_2)

Alkaline-Earths -- not as reactive as alkalis (two e^- to lose)

-- Ca and heavier ones react w/H_2O to form metal hydroxides

-- MgO is a protective oxide coating around substrate Mg

Hydrogen -- a nonmetal, but belongs to no family

-- reacts with other nonmetals to form molecular compounds

Halogens -- At isn't considered to be a halogen; little is known about it

-- at 25°C, F_2 and Cl_2 are gases, Br_2 is a liquid, I_2 is a solid

-- their exo. reactivity is dominated by their tendency to gain e^-

-- Cl_2 is added to water; the HOCl produced acts as a disinfectant

-- HF(aq) = weak acid; HCl(aq), HBr(aq), HI(aq) = strong acids

Noble Gases -- all are monatomic; have completely-filled *s* and *p*-orbitals

-- He, Ne, and Ar have no known compounds; Rn is radioactive

-- Kr has one known compound (KrF_2); Xe has a few (XeF_2, XeF_4, XeF_6)

Ionic Radius

Cations are _____ than parent atoms; anions are _____ than parent atoms.

EX. Compare the size of Fe, Fe^{2+}, Fe^{3+}, Br, and Br^-.

Electronegativity

<u>electronegativity</u>:

Electronegativity increases going...

Most electronegative element is...